Elements of Life (EL)

Revision guide

Clarity, confidence, and top grades.

What this storyline covers

- Atomic structure, spectra and electron configurations
- Fusion reactions and isotopes
- Mass spectrometry and relative masses
- The periodic table and Group 2 chemistry
- Bonding and the shapes of molecules
- Equations and amount of substance (moles)
- Ions: formulae, charge density and tests
- Titrations and acid–base calculations

Required practicals in this storyline: **PAG1** (reacting masses & moles), **PAG2** (standard solutions & titrations), **PAG4** (ion tests), plus making salts.



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How to use this guide

Elements of Life is the foundation of the whole course — the ideas here (moles, bonding, the periodic table, atomic structure) are revisited in every later storyline. Master this one and the rest become far easier.

How it is assessed

OCR B is context-led, so EL content can appear in **any** of the three written papers, in any context:

- **Paper 1** — Fundamentals of chemistry (110 marks)
- **Paper 2** — Scientific literacy, including the pre-release Advance Notice article (100 marks)
- **Paper 3** — Practical skills in chemistry (60 marks)

How to revise it

- Learn the **definitions in the teal boxes** word-for-word — they are marks you can guarantee.
- Work every **worked example** yourself before reading the solution.
- Note the **exam tips** (gold) and **common pitfalls** (linen) — these are where marks are won and lost.
- Use the **spec checklist** on the last page to confirm you can do every point.

1 Formulae, equations and amount of substance

Spec references: EL(a), EL(b), EL(c), EL(d). Practicals: PAG1, PAG2.

Key terms EL(a)

Key definitions

- **Atomic (proton) number, Z** — the number of protons in the nucleus.
- **Mass number, A** — the total number of protons and neutrons in the nucleus.
- **Isotopes** — atoms of the same element (same Z) with different numbers of neutrons (different A).
- **Relative isotopic mass** — the mass of one atom of an isotope relative to $\frac{1}{12}$ the mass of one atom of carbon-12.
- **Relative atomic mass, A_r** — the weighted mean mass of an atom of an element relative to $\frac{1}{12}$ the mass of one atom of carbon-12.
- **Relative molecular / formula mass, M_r** — the weighted mean mass of a molecule (or formula unit) relative to $\frac{1}{12}$ the mass of one atom of carbon-12; found by adding the A_r of every atom in the formula.
- **Avogadro constant, N_A** = 6.02×10^{23} mol^{-1} — the number of particles in one mole.
- **Mole** — the amount of substance containing 6.02×10^{23} particles (the number of atoms in exactly 12 g of carbon-12).

The mole and calculations EL(b), EL(c)

$$n = m \div M$$

amount (mol) = mass (g) ÷ molar mass ($g\ mol^{-1}$)

$$n = c \times V$$

amount (mol) = concentration ($mol\ dm^{-3}$) × volume (dm^3) [divide cm^3 by 1000]

- **Empirical formula** — the simplest whole-number ratio of atoms of each element. Divide each mass (or %) by A_r, divide through by the smallest, then scale to whole numbers.
- **Molecular formula** — the actual number of atoms of each element. Molecular formula = (M_r ÷ empirical-formula mass) × empirical formula.
- **Percentage composition by mass** = (mass of that element in the formula ÷ M_r) × 100.
- **Percentage yield** = (actual yield ÷ theoretical yield) × 100.
- **Water of crystallisation** — find moles of anhydrous salt and moles of water, then the simplest ratio (e.g. $MgSO_4 \cdot 7H_2O$).
- **Diluting a solution** — moles are unchanged, so $c_1V_1 = c_2V_2$.

Worked example — empirical and molecular formula

A compound contains 40.0% C, 6.7% H and 53.3% O by mass; M_r = 180. Find its molecular formula.

```
C: 40.0 / 12.0 = 3.33   H: 6.7 / 1.0 = 6.7   O: 53.3 / 16.0 = 3.33
Divide by smallest (3.33): C 1 : H 2 : O 1 → empirical formula CH2O (mass 30)
180 / 30 = 6 → molecular formula C6H12O6
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Worked example — water of crystallisation

Heating 4.96 g of hydrated $MgSO_4 \cdot xH_2O$ leaves 2.41 g of anhydrous $MgSO_4$ (M = 120.4). Find x.

$n(MgSO_4) = 2.41 / 120.4 = 0.0200$ mol

mass of water = $4.96 - 2.41 = 2.55$ g → $n(H_2O) = 2.55 / 18.0 = 0.142$ mol

ratio $H_2O : MgSO_4 = 0.142 / 0.0200 = 7.1$ → $x = 7$, so $MgSO_4 \cdot 7H_2O$

Concentration and titrations EL(c)

Key definition

A **standard solution** is a solution of accurately known concentration.

Worked example — titration calculation

25.0 cm^3 of sodium hydroxide is exactly neutralised by 23.50 cm^3 of 0.100 mol dm^{-3} hydrochloric acid. Find [NaOH].

$NaOH + HCl \rightarrow NaCl + H_2O$ (1 : 1)

$n(HCl) = 0.100 \times 23.50/1000 = 2.35 \times 10^{-3}$ mol

$n(NaOH) = 2.35 \times 10^{-3}$ mol (1 : 1)

$[NaOH] = (2.35 \times 10^{-3}) / (25.0/1000) = 0.094$ mol dm^{-3}

Practical techniques EL(b)(ii), EL(c)(ii) — PAG1, PAG2

- **Measuring mass of a solid:** weigh by difference (weigh the container + solid, transfer, then reweigh the container) so any residue is accounted for.
- **Making a standard solution:** weigh the solid accurately, dissolve in a beaker with a little distilled water, transfer with all washings into a volumetric flask, make up to the graduation mark (bottom of the meniscus on the line), stopper and invert to mix.
- **Titration:** rinse the burette with the solution it will hold; pipette a known volume into the conical flask (rinse the pipette with that solution); add indicator; run in from the burette, swirling, to the end point. Do a rough titre, then repeat for results **concordant within 0.10 cm^3** and use the mean of the concordant titres.
- **Diluting accurately:** pipette the aliquot into a volumetric flask and make up to the mark.

Balanced and ionic equations EL(d)

- Balance atoms on each side and include **state symbols** (s), (l), (g), (aq).
- **Ionic equation:** write soluble ionic species as separate ions and cancel the **spectator ions** (those unchanged on both sides). Example: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$.

Common pitfalls

- Forgetting to convert cm^3 to dm^3 (÷ 1000) in n = cV.
- Averaging **all** titres instead of only the concordant ones.
- Treating A_r as a simple average — it is a **weighted** mean.
- Leaving out state symbols when the question awards a mark for them.

2 Atomic structure

Spec references: EL(e), EL(f), EL(g), EL(h).

Orbitals and electron configuration EL(e), EL(f)

- An **orbital** is a region around the nucleus that can hold up to **2 electrons** (with opposite spins).
- **s-orbitals** are spherical; **p-orbitals** are dumbbell-shaped (three p-orbitals at right angles).
- Sub-shells: **s** (1 orbital, 2 e^-), **p** (3 orbitals, 6 e^-), **d** (5 orbitals, 10 e^-).
- **Filling order:** 1s, 2s, 2p, 3s, 3p, **4s, 3d**, 4p… (4s fills before 3d).
- **'Electrons in boxes':** each box is an orbital; fill orbitals singly before pairing, and show paired electrons with opposite-spin arrows (↑↓).

Species	Electron configuration	Note
Cl (Z = 17)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$	p-block atom
Cl^-	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ = [Ar]	gain 1 e^-
Ca^{2+}	[Ar]	lose the two 4s electrons
Cr (exception)	[Ar] $3d^5\ 4s^1$	half-filled 3d is stable
Cu (exception)	[Ar] $3d^{10}\ 4s^1$	full 3d is stable

Exam tip

When an atom forms a positive ion, remove electrons from **4s before 3d** (e.g. Fe [Ar]$3d^6 4s^2$ → Fe^{2+} [Ar]$3d^6$). See also DM(h) for d-block ions.

Development of the atomic model EL(g)

- Models became gradually more sophisticated: indivisible atom → 'plum pudding' → **nuclear model** → electron shells → orbitals.
- **Geiger–Marsden (Rutherford) experiment:** most α-particles passed straight through the gold foil, but a few were deflected or bounced back → evidence for a **small, dense, positively charged nucleus** with mostly empty space.
- Atoms and ions are described in terms of **protons, neutrons and electrons**.
- Evidence for electron shells comes from **successive ionisation enthalpies** (large jumps between shells, EL(q)) and **atomic spectra** (discrete lines, EL(w)).

Fusion reactions EL(h)

Key definition

Nuclear fusion — lighter nuclei join to form heavier nuclei under conditions of very high temperature and pressure (which overcome the repulsion between positively charged nuclei). This is how many elements are formed in stars.

$$^{2}_{1}H + {}^{3}_{1}H \rightarrow {}^{4}_{2}He + {}^{1}_{0}n$$

balance the top numbers (mass) and the bottom numbers (atomic) on each side

3 Bonding and structure

Spec references: EL(i), EL(j), EL(k), EL(l).

Types of bonding EL(i)

Key definitions

- **Ionic bond** — the electrostatic attraction between oppositely charged ions.
- **Covalent bond** — a shared pair of electrons; the electrostatic attraction between the shared pair and the two nuclei. There is a balance between repulsion (nucleus–nucleus, electron–electron) and attraction (nuclei–shared electrons).
- **Dative (coordinate) covalent bond** — a shared pair in which **both** electrons come from one atom (e.g. NH_4^+, where the lone pair on N is donated to H^+).

Use **dot-and-cross diagrams** to show the electron arrangement in ions, and in covalent and dative covalent bonds.

Structures and their properties EL(j), EL(l)

Structure	Bonding	Melting point	Conductivity	Solubility in water
Giant metallic	positive ions in a sea of delocalised electrons	high	conducts (delocalised electrons move)	insoluble
Giant ionic	lattice of + and − ions	high	only when molten / aqueous (ions free to move)	often soluble
Giant covalent	network of covalent bonds	very high	poor (graphite is the exception)	insoluble
Simple molecular	molecules held by weak intermolecular forces	low	poor	varies

Exam tip — the two property explanations OCR wants

- **Low melting point of molecular substances:** the electrostatic attractions *between molecules* are weaker than the electrostatic attractions in giant structures (only weak intermolecular forces are broken, not covalent bonds).
- **Conductivity:** requires charged particles that are free to move — delocalised electrons in metals, or ions in molten or aqueous ionic substances.

Sodium chloride lattice (EL(l)): a giant ionic lattice in which each ion is surrounded by six oppositely charged ions. Ionic bonding is the overall attraction in the lattice — attraction between ions of opposite charge and repulsion between ions of the same charge.

Shapes of molecules and ions EL(k)

Electron-pair repulsion

Electron pairs around the central atom repel one another and move as far apart as possible. **Lone pairs repel more strongly than bonding pairs** (lone–lone > lone–bonding > bonding–bonding), reducing the bond angle by about 2.5° per lone pair.

Bonding pairs	Lone pairs	Shape	Bond angle	Example
2	0	linear	180°	$BeCl_2$
3	0	trigonal planar	120°	BF_3
4	0	tetrahedral	109.5°	CH_4, NH_4^+
3	1	trigonal pyramidal	107°	NH_3
2	2	bent (non-linear)	104.5°	H_2O
6	0	octahedral	90°	SF_6

No treatment of hybridisation or molecular orbitals is required. Note the lone-pair effect: CH_4 109.5°, NH_3 107°, H_2O 104.5°.

4 Inorganic chemistry and the periodic table

Spec references: EL(m)–EL(s). Practical: PAG4.

The periodic table EL(m), EL(n)

- Elements are listed in order of **atomic (proton) number** and grouped by common properties (same outer-shell configuration). Use given data to predict the properties of an element in a group.
- Blocks: **s-block** (Groups 1–2), **p-block** (Groups 13–18), **d-block** (transition metals).
- **Melting-point trend across Periods 2–3:** giant metallic (Na→Mg→Al) rises (more delocalised electrons, stronger metallic bonding); giant covalent (C, Si) is very high; then simple molecular (P_4, S_8, Cl_2) is low (weak intermolecular forces), and the noble gas is lowest. Always explain in terms of **structure and bonding**.

Ions, names and formulae EL(o)

- Charge follows position: Group 1 → 1+, Group 2 → 2+, Al → 3+; Group 16 → 2−, Group 17 → 1−.
- Learn these ions: NO_3^-, SO_4^{2-}, CO_3^{2-}, OH^-, NH_4^+, HCO_3^-, Cu^{2+}, Zn^{2+}, Pb^{2+}, Fe^{2+}, Fe^{3+}.
- Without an oxidation state, **'nitrate' means NO_3^-** and **'sulfate' means SO_4^{2-}**.
- Build a formula by balancing total positive and negative charge.

Group 2 chemistry EL(p), EL(r)

Property of Mg → Ba	Trend down the group	Key equations / detail
Reaction with water	more vigorous	$M + 2H_2O \rightarrow M(OH)_2 + H_2$ (Mg reacts only slowly with cold water; with steam $Mg + H_2O \rightarrow MgO + H_2$)
Reaction with oxygen	more vigorous	$2M + O_2 \rightarrow 2MO$
Thermal stability of carbonate	increases (more stable)	$MCO_3 \rightarrow MO + CO_2$
Solubility of hydroxide	increases (more alkaline)	$Mg(OH)_2$ sparingly soluble; $Ba(OH)_2$ fairly soluble
Solubility of carbonate	decreases (all low)	Group 2 carbonates are insoluble

Charge density — EL(r)

Charge density = charge ÷ size of ion. Smaller ions of the same charge have higher charge density. A higher-charge-density cation **polarises (distorts) the carbonate ion** more, weakening it, so the carbonate decomposes at a lower temperature. Down Group 2 the cation gets larger, charge density falls, and the carbonates become **more thermally stable**.

Ionisation enthalpy EL(q)

Key definition

The **first ionisation enthalpy** is the enthalpy change when one electron is removed from each atom in one mole of gaseous atoms, forming one mole of gaseous 1+ ions.

$$M(g) \rightarrow M^+(g) + e^-$$

- **Across a period:** first ionisation enthalpy increases — the outermost electrons are in the same shell but are attracted by more protons (the small drops mid-period need not be explained).
- **Down a group:** first ionisation enthalpy decreases — the outer electrons are in shells further from the nucleus, with more shielding, so they are attracted less.
- Therefore s- and p-block metals become **more reactive down a group** (electrons lost more easily).

Solubility, precipitates and ion tests EL(s) — PAG4

Test for...	Reagent	Positive result
Carbonate, CO_3^{2-}	add dilute acid	effervescence; the gas turns limewater milky (CO_2)
Sulfate, SO_4^{2-}	acidify, then add Ba^{2+} ($BaCl_2$)	white precipitate ($BaSO_4$)
Chloride, Cl^-	acidify with HNO_3, add $AgNO_3$	white precipitate (AgCl)
Bromide, Br^-	acidify with HNO_3, add $AgNO_3$	cream precipitate (AgBr)
Iodide, I^-	acidify with HNO_3, add $AgNO_3$	yellow precipitate (AgI)
Cu^{2+}	add NaOH(aq)	pale blue precipitate, $Cu(OH)_2$
Fe^{2+}	add NaOH(aq)	green precipitate, $Fe(OH)_2$
Fe^{3+}	add NaOH(aq)	orange-brown precipitate, $Fe(OH)_3$

Exam tip — sequence of tests

Test for the anion and the cation in a logical order, and acidify before a precipitation test to remove interfering ions (for example, remove carbonate with acid before testing for sulfate or halide). Knowledge of the reaction of 3+ cations with CO_3^{2-} is not required.

5 Acid–base equilibria

Spec references: EL(t), EL(u). Practical: making salts.

Acids, bases and making salts EL(t)

Key definitions

- **Acid** — a proton (H^+) donor.
- **Base** — a proton acceptor.
- **Alkali** — a soluble base that releases OH^- ions in water.
- **Neutralisation** — acid + base → salt + water ($H^+ + OH^- \rightarrow H_2O$).

- Learn the mineral acids: **HCl, HNO_3, H_2SO_4**.
- **Soluble salt** (acid + alkali): titrate to find the volume, then repeat without indicator and crystallise. With an insoluble base or carbonate: add it in excess, filter off the excess, then crystallise.
- **Insoluble salt:** mix two soluble solutions (precipitation), then filter, wash and dry the precipitate.

Basic nature of Group 2 oxides and hydroxides EL(u)

- Group 2 oxides and hydroxides are **basic**: they react with water to give alkaline hydroxide solutions, and with acids to give a salt

and water (description and equations only).

- $MO + H_2O \rightarrow M(OH)_2$ $M(OH)_2 + 2HCl \rightarrow MCl_2 + 2H_2O$

6 Energy and matter

Spec references: EL(v), EL(w).

The electromagnetic spectrum EL(v)

In order of **increasing frequency and energy** (and decreasing wavelength): **infrared → visible → ultraviolet**.

Atomic spectra EL(w)

- Electron energy levels are **quantised**, so only fixed energy gaps are possible — this is why spectra are **line** spectra.
- **Emission spectrum:** electrons fall to lower levels and emit photons of fixed frequency → **bright coloured lines on a dark background**.
- **Absorption spectrum:** electrons absorb photons of fixed frequency and jump to higher levels → **dark lines on a bright/coloured background**.
- **Similarities:** both are line spectra; the lines are in the same positions for a given element; the lines get closer together (converge) at higher frequency; each series represents transitions to or from a particular energy level.

$$\Delta E = h\nu$$

energy of a transition (J) = Planck constant (6.63×10^{-34} J s) × frequency (Hz)

$$c = \nu\lambda$$

speed of light (3.00×10^{8} m s^{-1}) = frequency (Hz) × wavelength (m)

Worked example — energy of a spectral line

A line in the hydrogen spectrum has wavelength 656 nm. Find its frequency and the energy of the transition.

$\nu = c / \lambda = (3.00 \times 10^{8}) / (656 \times 10^{-9}) = 4.57 \times 10^{14}$ Hz

$\Delta E = h\nu = (6.63 \times 10^{-34}) \times (4.57 \times 10^{14}) = 3.03 \times 10^{-19}$ J

Flame colours EL(w)(v)

Heat in a flame excites electrons to higher levels; as they fall back they emit visible light of a characteristic frequency.

Ion	Li^+	Na^+	K^+	Ca^{2+}	Ba^{2+}	Cu^{2+}
Flame colour	red	yellow	lilac	brick-red	green	blue-green

7 Modern analytical techniques

Spec reference: EL(x).

Mass spectrometry and relative atomic mass EL(x)

A mass spectrum gives the mass and relative abundance of each isotope. Use these to calculate A_r:

$$A_r = \Sigma(\text{isotopic mass} \times \%\ \text{abundance}) \div 100$$

Worked example — A_r of chlorine

Chlorine's mass spectrum shows ^{35}Cl at 75.0% and ^{37}Cl at 25.0%.

$A_r = (35 \times 75.0 + 37 \times 25.0) / 100 = (2625 + 925) / 100 = 35.5$

P Required practicals in this storyline

These PAG activities are most likely to be examined in Paper 3, but the techniques can appear in any paper.

PAG	Activity	What to be able to do
PAG1	Reacting masses and moles	weigh by difference; determine a formula (e.g. water of crystallisation by heating to constant mass); calculate percentage yield and percentage composition
PAG2	Standard solutions & titrations	make up a standard solution in a volumetric flask; dilute accurately; carry out an acid–base titration to concordant results and calculate concentration
PAG4	Ion tests	carry out precipitation reactions of the EL(s) ions and use a logical sequence of tests to identify an unknown salt
—	Making salts	prepare a soluble salt (acid + base/carbonate, then crystallise) or an insoluble salt (precipitation, then filter and dry); calculate percentage yield

Exam tip — practical accuracy language

Marks for practical questions usually come from precise technique: weighing **by difference**, reading the bottom of the **meniscus**, transferring with **washings**, repeating to **constant mass** or **concordant titres**, and identifying the main source of **uncertainty** (the apparatus with the largest percentage error).

Spec coverage checklist — can you do every point?

Every assessable learning outcome for Elements of Life, with where it is covered in this guide. Tick each one once you can do it confidently.

EL(a)	Define Z, A, isotope, N_A, relative isotopic / atomic / formula / molecular mass — §1
EL(b)	Moles calculations: masses, empirical & molecular formulae, % composition, % yield, water of crystallisation; measure mass of solids — §1
EL(c)	Moles calculations: concentration, titration and dilution; measure volumes and prepare a standard solution — §1
EL(d)	Balanced full and ionic equations with state symbols — §1
EL(e)	Represent electrons in orbitals; shapes of s- and p-orbitals — §2
EL(f)	Electron configurations of atoms H–Kr, s/p-block ions, and outer sub-shells of other periods — §2
EL(g)	Development of the atomic model; Geiger–Marsden evidence; particles in atoms/ions; evidence for shells — §2
EL(h)	Fusion reactions and nuclear equations — §2
EL(i)	Bonding as electrostatic forces; dot-and-cross for ions, covalent and dative bonds — §3
EL(j)	Giant (metallic, ionic, covalent) and simple molecular structures and their properties — §3
EL(k)	Electron-pair repulsion: shapes and bond angles of molecules and ions — §3
EL(l)	Structure of a sodium chloride type lattice — §3
EL(m)	The periodic table; predicting properties; s-, p- and d-blocks — §4
EL(n)	Melting-point trends across Periods 2 and 3 — §4
EL(o)	Charge from position; names and formulae of the listed ions and their compounds — §4
EL(p)	Group 2: reactions with water and oxygen; carbonate stability; hydroxide and carbonate solubility — §4
EL(q)	Ionisation enthalpy: define, equation, trends across/down, link to reactivity — §4
EL(r)	Charge density and the thermal stability of Group 2 carbonates — §4
EL(s)	Solubility, precipitate colours and ion tests; a sequence of tests to identify a salt — §4
EL(t)	Acid, base, alkali, neutralisation; making soluble and insoluble salts; mineral acids — §5
EL(u)	Basic nature of Group 2 oxides and hydroxides — §5
EL(v)	The electromagnetic spectrum: infrared, visible, ultraviolet — §6
EL(w)	Atomic spectra; absorption/emission; $\Delta E = h\nu$; $c = \nu\lambda$; flame colours — §6
EL(x)	Use mass-spectrum data to find isotopic abundance and A_r — §7

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